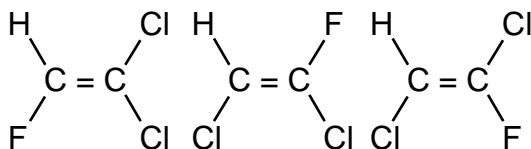


(b) Why is the PH_3 dipole moment less than that for NH_3 ?

The molecular structures of PH_3 and NH_3 are the same: trigonal pyramidal. But the P–H bond dipole moment is smaller than the N–H bond dipole moment because P is less electronegative than N.

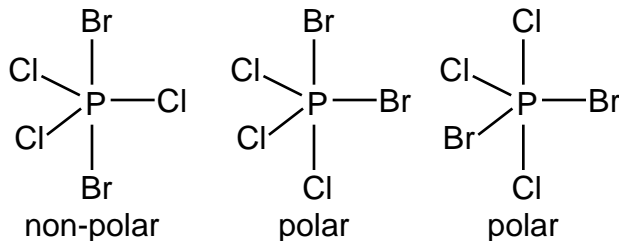
(c) There are **3** distinctly different molecules with the empirical formula C_2HFCl_2 .

And here they are:



(d) The molecule PBr_2Cl_3 **could have** a dipole moment.

There are three isomers, one non-polar, and two polar:



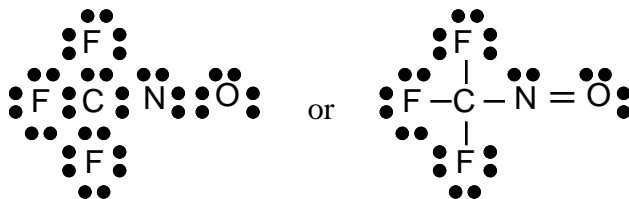
(e) The molecular orbital electron configuration of the carbide ion, C_2^{2-} , is $\sigma_{1s}^2 \sigma_{1s}^{*2} \sigma_{2s}^2 \sigma_{2s}^{*2} \pi_{2p}^4 \sigma_{2p_z}^2$.

C_2^{2-} is isoelectronic to N_2 , with a triple bond: $\text{C}\equiv\text{C}^{2-}$.

3. The molecule CF_3NO is a lovely pale blue gas that decomposes when irradiated according to the net reaction



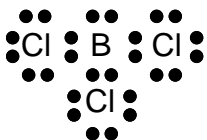
forming hexafluoroethane, CF_3CF_3 , and nitric oxide, NO . Draw the Lewis electron dot structure of CF_3NO , and then answer the following questions.



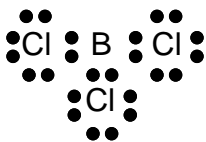
- (a) CF_3NO has **one** π bond. (It's in the $\text{N}=\text{O}$ double bond.)
 (b) The hybridization scheme used by the N atom is **sp^2** .
 (c) The hybridization scheme used by the C atom is **sp^3** .
 (d) The hybridization scheme used by the O atom is **sp^2** .
 (e) The molecule has **12** lone pair(s) of electrons.
 (f) The C–N–O bond is **not** linear. (because N is sp^2 hybridized)
 (g) The formal charge on the N atom is **zero**. ($5 \text{ valence } e^- \text{'s} - 2 \text{ nonbonded electrons} - (6/2) \text{ bonded electrons} = 0$)
 (h) The strongest bond in CF_3NO is most likely between the atoms **N and O**. (double bond)
 (i) The following are paramagnetic: **NO** (Odd number of electrons)

4. Each of the following Lewis dot structures has one or more deficiency that makes it somewhere between less than optimal or dead wrong. Draw the optimum structure and state what is bad about the structures below.

(a) BCl_3



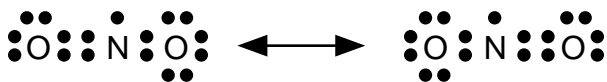
This is dead wrong—too many electrons! Boron has only three valence electrons, making BCl_3 an electron-deficient molecule:



(b) NO_2



There are 9 e^- 's on N. Optimum (resonance, with the "odd" electron on the less electronegative N atom):

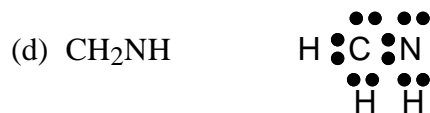


(c) OCS

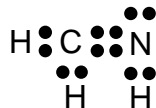


Bad formal charges on O (-1) and S ($+1$) compared to optimal structure with no formal charges:





No octet around N. Optimal:



5. Some dipole moment questions:

(a) The dipole moment of ClF is 0.8881 D = 2.96×10^{-30} C m, and the bond is 11.3% ionic. What is the ClF bond length?

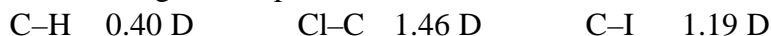
We use the expression relating the dipole moment, μ , to the bond length, R , and the fractional ionicity δ (11.3% ionic means $\delta = 0.113$):

$$R/\text{\AA} = \frac{(\mu/\text{D})(0.2082)}{\delta} = \frac{(0.8881)(0.2082)}{0.113} = 1.64 \quad \text{or } R = 1.64 \text{\AA}$$

Equivalently, we can express these quantities in their fundamental SI units:

$$R = \frac{\mu}{\delta e} = \frac{2.96 \times 10^{-30} \text{ C m}}{(0.113)(1.602 \times 10^{-19} \text{ C})} = 1.64 \times 10^{-10} \text{ m} \quad \text{or } R = 1.64 \text{\AA}$$

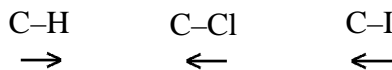
(b) Given the following bond dipole moments:



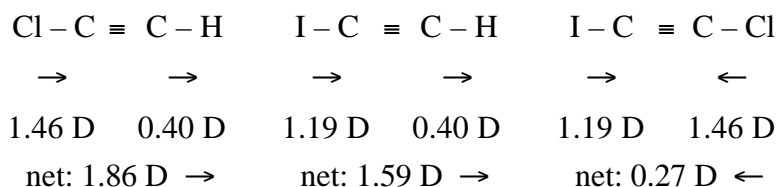
predict the magnitude and direction for the dipole moments of the linear molecules:

- (i) Cl–C≡C–H
 (ii) I–C≡C–H
 (iii) I–C≡C–Cl

First, we use electronegativity values (which will be provided on the exam if needed) to deduce the direction of each of these bond dipole moments. The relevant electronegativities are: C, 2.55; H, 2.20; Cl, 3.16; and I, 2.66. Using these values and remembering that the dipole moment vector points from the negative (more electronegative) to the positive end of the dipole, we can write



This lets us write bond dipole vectors in the three molecules of interest and combine them to form the total dipole moment magnitudes and directions:

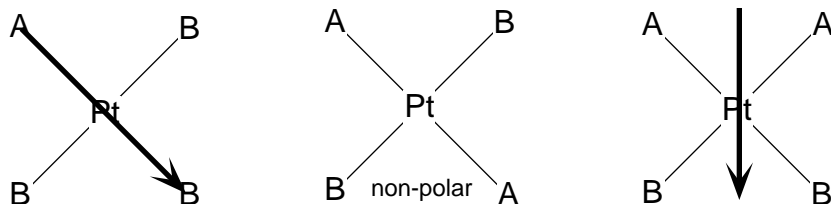


(c) The molecule Cl_2CO , phosgene, has a net dipole moment that points one way or another along one bond in the molecule. Which bond is it, and how can you tell? (Don't worry about its direction along that bond.)

Cl_2CO is a planar molecule (like formaldehyde) with the central C atom double-bonded to the O and single-bonded to the two Cl atoms so that the C=O bond is a symmetry axis. The dipole points along this bond, because the Cl–C bond moment components perpendicular to this direction cancel each other, while their components along this direction combine with the C=O bond moment to produce a net moment that is non-zero.

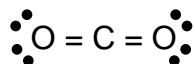
(d) There are many compounds that contain Pt with four things bonded to it in a square planar arrangement. (The chemotherapy drug called cisplatin is one such.) Consider the two generic types of compounds with empirical formulas PtAB_3 and PtA_2B_2 in which A and B are atoms bonded to the Pt. Let's assume A is more electronegative than B. Draw all possible isomers of these compounds, and indicate the direction of the net dipole moment (if there is one).

There is only one PtAB_3 isomer, but two PtA_2B_2 isomers, one of which is non-polar:

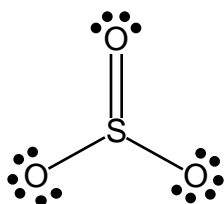


6. Give the electron pair geometry and the molecular geometry of each of the following oxides. Draw molecular structures for each, showing all single bonds, multiple bonds, and lone pairs to support your geometry assignments.

(a) CO_2 electron pair geometry: *linear* molecular geometry: *linear*



(b) SO_3 electron pair geometry: *trigonal planar*
molecular geometry: *trigonal planar*



(plus two resonance forms in which this structure is rotated 120° first clockwise and then counter-clockwise)

- (c) XeO_4 electron pair geometry: *tetrahedral*
molecular geometry: *tetrahedral*

